1. A sample of gas has an initial volume of 158 mL at a pressure of 735 mm Hg and a temperature of 34 °C. If the gas is compressed to a volume of 108 mL and heated to a temperature of 85 °C, what is its final pressure in millimeters of mercury?

\[ P_1 = 735 \text{ mm Hg} \]
\[ T_1 = 34 \degree \text{C} = 307.15 \text{ K} \]
\[ V_1 = 158 \text{ mL} = 0.158 \ell \]
\[ V_2 = 108 \text{ mL} = 0.108 \ell \]

\[ n_1 = n_2 \]

\[ \frac{P_1 \cdot V_1}{T_1} = \frac{P_2 \cdot V_2}{T_2} \quad \text{(Combined gas law)} \]

\[ P_2 = \frac{735 \text{ mm Hg} \times \frac{1 \text{ atm}}{760 \text{ mm Hg}} \times 0.158 \ell \times 358.15 \text{ K}}{0.108 \ell \times 307.15 \text{ K}} \]

\[ P_2 = 1.65 \text{ atm} = 1.25 \times 10^3 \text{ mm Hg} \]

\( (3SF) \)
2. A 4.8-L sample of helium gas contains 0.22 mol of helium. How many additional moles of helium gas must be added to the sample to obtain a volume of 6.4 L? Assume constant temperature and pressure.

\[ V_1 = 4.8 \text{ L} \quad V_2 = 6.4 \text{ L} \]

\[ n_1 = 0.22 \text{ mol} \quad n_2 = ? \]

\[ \frac{V_1}{n_1} = \frac{V_2}{n_2} \quad \text{(Avogadro's Law)} \]

\[ n_2 = \frac{6.4 \text{ L} \times 0.22 \text{ mol}}{4.8 \text{ L}} = 0.29 \text{ mol} \]

\[ \text{mol to add} = 0.29 - 0.22 = 0.07 \text{ mol}. \]
3. A sample of gas has a mass of 0.311 g. Its volume is 0.225 L at a temperature of 55 °C and a pressure of 886 mm Hg. Find its molar mass.

\[ m = 0.311 \text{ g} \]
\[ V = 0.225 \text{ L} \]
\[ T = 55 ^\circ \text{C} + 273.15 = 328.15 \text{ K} \]
\[ P = 886 \text{ mm Hg} \]

molar mass (g/mol) = ?

P. \ V = n \cdot R \cdot T

P. \ V = \frac{\text{mass}}{\text{molar mass}} \cdot R \cdot T

P = 886 \text{ mm Hg} \times \frac{1 \text{ atm}}{760 \text{ mm Hg}} = 1.1658 \text{ atm} \cdot (3 \text{sf})

molar mass = \frac{0.311 \text{ g} \times 0.082058 \text{ atm} \cdot \text{L/mol} \cdot \text{K}}{1.1658 \text{ atm} \times 0.225 \text{ L}}

molar mass = 31.9 \text{ g/mol} \ (3 \text{sf})
4. How many liters of oxygen gas form when 294 g of KClO₃ completely react in this reaction (which is used in the ignition of fireworks)?

\[ 2 \text{KClO}_3(s) \rightarrow 2 \text{KCl(s)} + 3 \text{O}_2(g) \]

Assume that the oxygen gas is collected at \(P = 755 \text{ mm Hg}\) and \(T = 305 \text{ K}\).

Gas in chemical reactions:

\[ M_{\text{KClO}_3} = 294 \text{ g} \]

\[ P = 755 \text{ mm Hg} \] (of oxygen gas)

\[ T = 305 \text{ K} \]

\[ V_{\text{O}_2} = ? \text{ L} \]

Let \( n_{\text{O}_2} \) be the number of moles \( \text{O}_2 \) produced.

\[ n_{\text{O}_2} = \frac{0.5 \text{ mol KClO}_3}{122.5 \text{ g KClO}_3} \times \frac{3 \text{ mol O}_2}{2 \text{ mol KClO}_3} \]

\[ n_{\text{O}_2} = 3.60 \text{ mol} \text{ O}_2 \] (3 SF)

\[ P \cdot V = n \cdot R \cdot T \]

\[ V_{\text{O}_2} = \frac{3.60 \text{ mol} \times 0.082058 \text{ L/ mol k} \times 305 \text{ K}}{755 \text{ mm Hg} \times \frac{1 \text{ atm}}{760 \text{ mm Hg}}} \]

\[ V_{\text{O}_2} = 90.7 \text{ L} \] (3 SF)
5. In an experiment, hydrogen gas is produced by the reaction of HCl with Zn and it is collected over water. If the volume of the collected gas is 192 mL at 19°C and 769 mmHg total pressure, what is the mass of collected hydrogen in g?

\[ P_{\text{H}_2} = P_{\text{total}} - P_{\text{H}_2O@19°C} = (769 - 16.5) \text{ mmHg} = 752.5 \text{ mmHg} \]

Now we can calculate the amount of hydrogen collected using the ideal gas equation:

\[ n_{\text{H}_2} = \frac{PV}{RT} = \frac{752.5 \text{ mmHg} \times 0.192 \text{ L}}{760 \text{ mmHg} \times 0.082058 \text{ atm.L/mol.k} \times (273.15+19) \text{ K}} \]

\[ n_{\text{H}_2} = 0.0077703 \text{ mol} \]

If \( m \) is the mass of hydrogen in grams,

\[ m = n_{\text{H}_2} \times \text{MM}_{\text{H}_2} = 0.0077703 \text{ mol} \times 2.01588 \text{ g/mol} \]

\[ = 0.0157 \text{ g (3SF)} \]
6. Nickel forms a gaseous compound of the formula Ni(CO)$_x$. What is the value of $x$ given the fact that under the same conditions of temperature and pressure, methane (CH$_4$) effuses 3.3 times faster than the compound?

\[
\frac{r_{\text{CH}_4}}{r_{\text{Ni(CO)}_x}} = \sqrt{\frac{M_{\text{Ni(CO)}_x}}{M_{\text{CH}_4}}} \quad \text{Graham's Law of Diffusion}
\]

\[
\frac{3.3}{1.0} = \sqrt{\frac{M_{\text{Ni(CO)}_x}}{16.04 \text{ g/mol}}}
\]

\[
M_{\text{Ni(CO)}_x} = 174.7 \text{ g/mol}
\]

To find the value of $x$, we first subtract the molar mass of Ni from 174.7 g/mol.

\[
174.7 \text{ g} - 58.69 \text{ g} = 116.0 \text{ g}.
\]

116.0 g is the mass of CO in 1 mole of the compound. The mass of 1 mole of CO is 28.01 g.

\[
x = \frac{116.0 \text{ g}}{28.01 \text{ g}} = 4.141 \approx 4
\]

Therefore, the value of $x$ is 4. Ni(CO)$_4$
7. Using the data shown in Table 5.4, calculate the pressure exerted by 2.50 moles of CO₂ confined in a volume of 5.00 L at 450 K. Compare the pressure with that predicted by the ideal gas equation.

Here, we compare the pressure as determined by the van der Waals' equation and the ideal gas equation:

**van der Waals' equation:**

\[
(P + \frac{a n^2}{V^2})(V - nb) = nRT
\]

For CO₂: \(a = 35.9 \text{ atm}. \text{L}^2/\text{mol}^2\) \(b = 0.0427 \text{ L}/\text{mol}\)

\[
P = \frac{nRT}{(V - nb) - \frac{an^2}{V^2}}
\]

\[
P = \frac{(2.50 \text{ mol}) (0.082058 \frac{\text{atm}. \text{L}}{\text{mol}. \text{K}}) (450 \text{ K})}{\left[ (5.00 \text{ L}) - (2.50 \text{ mol} \times 0.0427 \text{ L/mol}) \right]}
\]

\[
P = 18.0 \text{ atm.}
\]

**Ideal gas equation:**

\[
P = \frac{nRT}{V} = \frac{(2.50 \text{ mol}) (0.082058 \frac{\text{atm}. \text{L}}{\text{mol}. \text{K}}) (450 \text{ K})}{5.00 \text{ L}}
\]

\[
P = 18.5 \text{ atm.}
\]

Since the pressure calculated using van der Waals' equation is comparable to the pressure calculated using the ideal gas equation, we conclude that CO₂ behaves fairly ideally under these conditions.
8. Calculate the density of a gaseous organic compound at STP conditions that has a molar mass of 67.9 g/mol?

STP conditions:

\[ T = 0^\circ C = 273.15 K \]

\[ P = 1 \text{ atm} \]

\[ d = \frac{P \cdot M}{R \cdot T} = \frac{(1 \text{ atm})(67.9 \text{ g/mol})}{(0.082059 \text{ atm L/mol K})(273.15 \text{ K})} = 3.03 \text{ g/L} \]
9. Consider the following apparatus. Calculate the partial pressures of helium and neon after the stopcock is open. The temperature remains constant at 16°C.

\[ P = \frac{nRT}{V} \]

\[ n_{\text{He}} = \frac{PV}{RT} = \frac{(0.63 \text{ atm})(1.2 \text{ L})}{(0.082058 \text{ atm L/mol K})(16273.15 \text{ K})} = 0.032 \text{ mol He} \]

\[ n_{\text{Ne}} = \frac{PV}{RT} = \frac{(2.8 \text{ atm})(3.4 \text{ L})}{(0.082058 \text{ atm L/mol K})(16273.15 \text{ K})} = 0.040 \text{ mol Ne} \]

The total pressure is:

\[ P_{\text{tot}} = \frac{n_{\text{He}} + n_{\text{Ne}}}{V_{\text{tot}}} = \frac{(0.32+0.40)\text{ mol}(0.08205 \text{ L atm/mol K})(16273.15 \text{ K})}{(1.2 + 3.4) \text{ L}} = 2.2 \text{ atm} \]

The total pressures of He and Ne are:

\[ P_{\text{He}} = \frac{0.032 \text{ mol}}{(0.032 + 0.40)\text{ mol}} \times 2.2 \text{ atm} = 0.16 \text{ atm} \]

\[ P_{\text{Ne}} = \frac{0.40 \text{ mol}}{(0.032 + 0.40)\text{ mol}} \times 2.2 \text{ atm} = 2.0 \text{ atm} \]